Knight: Chapter 17

Work, Heat, & the 1^{st} Law of Thermodynamics

(*Thermal Properties of Matter, Calorimetry, & The Specific Heat of Gases*)

Quiz Question 1

Two liquids, A and B, have equal masses and equal initial temperatures. Each is heated for the same length of time over identical burners. Afterward, liquid A is hotter than liquid B. Which has the larger specific heat?

$$\Delta E_{th_A} = \Delta E_{th_B}$$

1. Liquid A.
2. Liquid B.
3. There' s not enough information to tell.

Thermal Properties of Matter

If the *temperature* of a system changes,
the *thermal energy* of the system changes by

$$\Delta E_{th} = Mc\Delta T = Q + W$$

$n = \frac{M}{M_{mol}}$

$M = n(M_{mol})$

- If NO work is done on the system, then the heat needed to bring about a temp change ΔT is:

$$\boxed{Q = Mc\Delta T}$$

$Q = n(M_{mol})c\Delta T$

where c is the *specific heat*.

- Specific heat can be thought of as the *thermal inertia* of a substance!

Thermal Properties of Matter

The *molar specific heat*…

- the amount of energy that raises the temp of 1 mol of a substance by 1K

$$\boxed{Q = nC\Delta T}$$

TABLE 17.2 Specific heats and molar specific heats of solids and liquids

Substance	c (J/kg K)	C (J/mol K)
Solids		
Aluminum	900	24.3
Copper	385	24.4
Iron	449	25.1
Gold	129	25.4
Lead	128	26.5
Ice	2090	37.6
Liquids		
Ethyl alcohol	2400	110.4
Mercury	140	28.1
Water	4190	75.4

Notice:

- Most elemental solids have
$$C \simeq 25\ \mathrm{J/mol\ K}$$
- Why?

Thermal Properties of Matter

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Notice:

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- Why? Answer in CH 18

Quiz Question 2

If you heat a substance in a rigid container, is it possible that the temperature of the substance remains unchanged?

$PV = nRT$

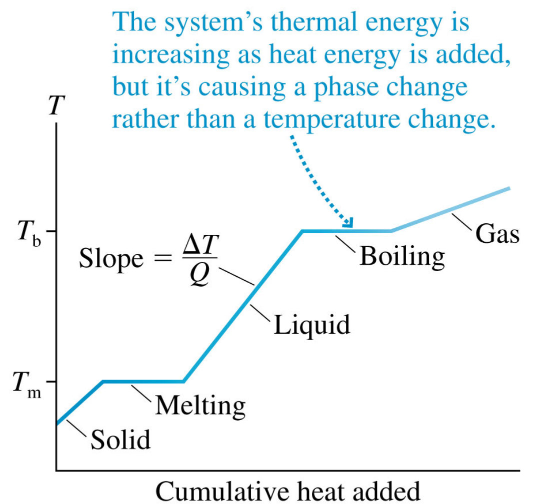
1. Yes. Only in a phase change
2. No.

Phase Change & Heat of Transformation

Suppose you start with a system in its solid phase and heat it at a steady rate…

- The heat needed for a phase change is:

$$\boxed{Q = \pm M L_{f,v}}$$



Notice:

- L_f is the *heat of fusion*, L_v is the *heat of vaporization*.
- Positive/negative sign must be included by hand!
- The ~~*heat of fusion/vaporization*~~ can be measured from the slope. Specific heat

Quiz Question 3

1 kg of silver (c = 234 J/kg K) is heated to 100°C. It is then dropped into 1 kg of water (c = 4190 J/kg K) at 0°C in an insulated beaker. After a short while, the common temperature of the water and silver is

1. 0°C.
2. between 0°C and 50°C.
3. 50°C.
4. between 50°C and 100°C.
5. 100°C.

Calorimetry

Consider 2 systems with different temps T_1 and T_2 that interact thermally with each other but are isolated from everything else…

- Heat will flow *from* the *hotter to* the *colder* system until they reach an equilibrium temp T_f.

$$Q_1 + Q_2 = 0$$

- In general:

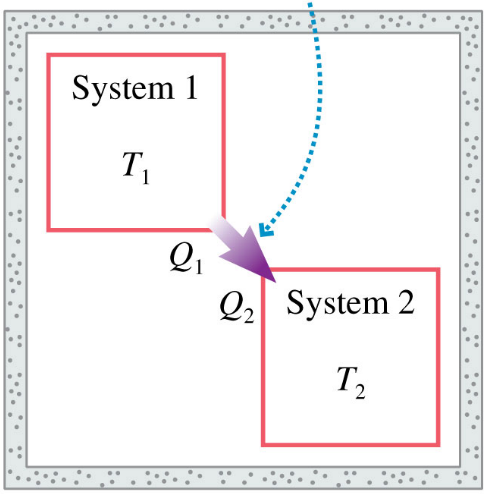
$$\boxed{Q_{net} = Q_1 + Q_2 + Q_3 + \ldots = 0}$$

Heat energy is transferred from system 1 to system 2. Energy conservation requires

$$|Q_1| = |Q_2|$$

Opposite signs mean that

$$Q_{net} = Q_1 + Q_2 = 0$$



Problem-Solving Strategy: Calorimetry Problems

PROBLEM-SOLVING STRATEGY 17.2 **Calorimetry problems**

MODEL Identify the interacting systems. Assume that they are isolated from the larger environment.

VISUALIZE List known information and identify what you need to find. Convert all quantities to SI units.

SOLVE The mathematical representation, which is a statement of energy conservation, is

$$Q_{net} = Q_1 + Q_2 + \cdots = 0$$

- For systems that undergo a temperature change, $Q = Mc(T_f - T_i)$. Be sure to have the temperatures T_i and T_f in the correct order.
- For systems that undergo a phase change, $Q = \pm ML$. Supply the correct sign by observing whether energy enters or leaves the system.
- Some systems may undergo a temperature change *and* a phase change. Treat the changes separately. The heat energy is $Q = Q_{\Delta T} + Q_{phase}$.

ASSESS Is the final temperature in the middle? T_f that is higher or lower than all initial temperatures is an indication that something is wrong, usually a sign error.

Exercise 15

i.e. 17.5: Calorimetry with a phase change

Your 500 mL soda is at 20° C, room temperature, so you add 100 g of ice from the -20° C freezer.

Does all the ice melt? If so, what is the final temperature? If not, what fraction of the ice melts? Assume you have a well-insulated cup.

$Q_{ice} + Q_{soda} = 0$

$V = 5.0 \times 10^{-4} m^3$

$m_{ice} c_{ice} (20K) + m_{ice} L_{f_{H_2O}} + m_{ice} c (T_F - 273K) + m_S c_{H_2O} (T_F - 293K)$

$T_S = 293 \text{ K} \quad T_I = 253 \text{ K}$

$m_I = 0.100 \text{ Kg}$

$c_{H_2O} = 4{,}186 \text{ J/KgK}$

$c_I = 2{,}090 \text{ J/KgK}$

$L_{H_2O} = 3.33 \times 10^5 \text{ J/Kg}$

$\rho_{H_2O} = \frac{m_{H_2O}}{V} = m_{H_2O} = \rho_{H_2O} V = (1{,}000 \text{ Kg/m}^3)(5.0 \times 10^{-4} \text{ m}^3) = 0.50 \text{ Kg}$

i.e. 17.6: Three interacting systems

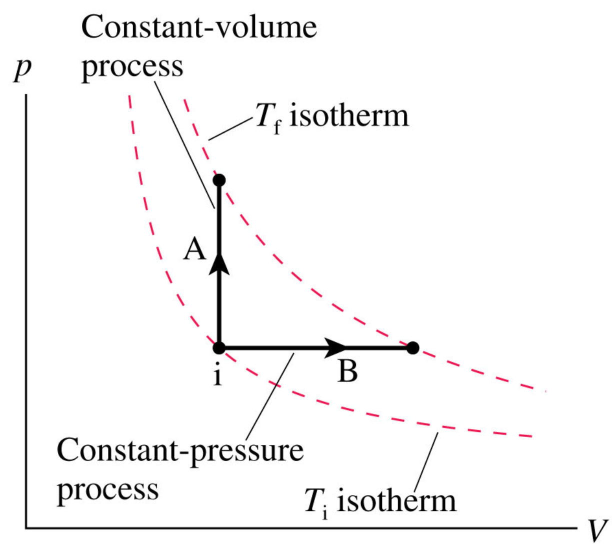
A 200 g piece of iron at 120° C and a 150 g piece of copper at -50° C are dropped into an insulated beaker containing 300 g of ethyl alcohol at 20° C.

What is the final temperature?

The Specific Heat of Gases

Consider the two processes A & B…

- Both have the *same* ΔT, therefore the *same* ΔE_{th}, but they require *different* amounts of Q.
  - The reason is that work is done in process B but not in process A.
- The total change in thermal energy for any process, due to work and heat, is:



$$\boxed{\Delta E_{th} = nC_V \Delta T} \quad \text{(any ideal-gas process)}$$

The Specific Heat of Gases

Define 2 different versions of the specific heat of gases…

- one for constant-*volume* processes.
- one for constant-*pressure* processes.

- The quantity of heat needed to change the temperature of n moles of gas by ΔT is:

$$\boxed{\begin{aligned} Q &= nC_V\Delta T \quad \text{(temp change at constant volume)} \\ Q &= nC_P\Delta T \quad \text{(temp change at constant pressure)} \end{aligned}}$$

- where C_V is the *molar specific heat at constant volume* & C_P is the *molar specific heat at constant pressure*.

The Specific Heat of Gases

What if neither the pressure nor the volume is constant?

The Specific Heat of Gases

What if neither the pressure nor the volume is constant?

- no direct way to relate Q to ΔT!
- Use the 1st law of thermodynamics:

$$Q = \Delta E_{th} - W$$